First steps!

Calculate the concentration of Cl^- ions if $PbCl_2$ contains 0.0145 $molL^{-1}$ of Pb^{2+} ions

$PbCl_2$

1 : 2

0.0145 : 2 x 0.0145

Answer: 0.029 $molL^{-1}$

Calculate the solubility product K_s if concentration of Pb^{2+} is 0.0145 $molL^{-1}$

$PbCl_2(s) \rightleftharpoons Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$

$Ks = [Pb^{2+}][Cl^-]^2$

$= (0.0145)(0.029)^2$

Answer: 1.22 x 10^{-5}

Ks, solubility product/constant this is an equilibrium constant which applies to saturated solutions

a large Ks value = high solubility

a low Ks value = low solubility/insoluble

Working out the solubility product, K_s

Given Ks = 1.81x 10^{-10} of AgCl, find the concentration of Ag^+ ions

$Ks = [Ag^+][Cl^-]$

$= x . x$

$= x^2$

$\sqrt{Ks} = x$

Answer: 1.35 x 10^{-5} mol L^{-1}

Given Ks = 7.1x 10^{-5} of $CaSO_4$, find the conc. of SO_4^{2-} ions in $CaSO_4$

$Ks = [Ca^{2+}][SO_4^{2-}]$

$Ks = x . x$

$\sqrt{Ks} = x$

$\sqrt{7.1 \times 10^{-5}} = x$

Answer: 8.4 x 10^{-3} mol L^{-1}

…more working out Ks

Given Ks = 6.6 x 10^{-6}, find the concentration of Cu^{2+} ions in $CuBr_2$

$Ks = [Cu^{2+}][Br^-][Br^-]$

$Ks = [Cu^{2+}][Br^-]^2$

$Ks = x . 2x^2$

$Ks = 4x^3$

$\sqrt[3]{\frac{Ks}{4}} = x$

$\sqrt[3]{1.65 \times 10^{-6}} = x$

Answer: 0.0118 $molL^{-1}$

…and more working out Ks

Given Ks = 1.2 x 10^{-5}, find the concentration of Ag^+ ions in Ag_2SO_4

$Ks = [Ag^+][Ag^+][SO_4^{2-}]$

$Ks = [Ag^+]^2[SO_4^{2-}]$

$Ks = 2x^2 . x$

$Ks = 4x^3$

$\sqrt[3]{\frac{Ks}{4}} = x$

$\sqrt[3]{3 \times 10^{-6}} = x$

So, x = 0.0171 $molL^{-1}$

Answer: the concentration of Ag^+ ions is twice x, therefore 0.0342$molL^{-1}$

the **"common ion"** effect: precipitation can occur ie solubility decreases if an ion is added to a solution that already contains that ion

Calculate the solubility of $Fe(OH)_2$ in a 0.05 $molL^{-1}$ solution of NaOH

Ks of $Fe(OH)_2$ = 7.9 x 10^{-16}

$Ks = [Fe^{2+}][OH^-]^2$

assume that $[OH^-] = 0.05molL^{-1}$

$7.9 \times 10^{-16} = [Fe^{2+}](0.05^-)^2$

$\frac{7.9 \times 10^{-16}}{(0.05)^2} = [Fe^{2+}]$

Answer: 3.16 x 10^{-13}

Calculate the solubility of AgCl in 0.1$molL^{-1}$ NaCl.

Ks of AgCl is 2 x 10^{-10}

$Ks = [Ag^+][Cl-]$

assume that [Cl-] = 0.1$molL^{-1}$

$2 \times 10^{-10} = [Ag^+](0.1)$

$\frac{2 \times 10^{-10}}{(0.1)} = [Ag^+]$

Answer: 2 x 10^{-9}

If a solution is not in equilibrium the term is **Ionic product (IP)**

IP > Ks a ppt will occur

IP = Ks a saturated solution

IP < Ks there is no ppt

Consider whether a precipitate will occur if 50mL of 0.02$molL^{-1}$ Na_2CO_3 is mixed with 50mL of 0.05$molL^{-1}$ $CaCO_3$.

Ks $(CaCO_3)$ = 3.4 x 10^{-9}

As the two solutions are mixed together the volume doubles so the concentration of all ions is halved

$[Ca^{2+}] = \frac{0.05}{2}$ $\quad [CO_3^{2-}] = \frac{0.05}{2} + \frac{0.02}{2}$

$= 0.025$ $\quad = 0.035$

$IP = [Ca^{2+}][CO_3^{2-}] = 0.025 \times 0.035 = 8.75 \times 10^{-4}$

Answer: IP > Ks so a precipitate **will** occur